

First steps!

Calculate the concentration of Cl^- ions if PbCl_2 contains $0.0145 \text{ mol L}^{-1}$ of Pb^{2+} ions

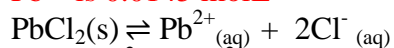


1 : 2

$$0.0145 : 2 \times 0.0145$$

Answer: 0.029 mol L^{-1}

Calculate the solubility product K_s if concentration of Pb^{2+} is $0.0145 \text{ mol L}^{-1}$



$$K_s = [\text{Pb}^{2+}] [\text{Cl}^-]^2$$

$$= (0.0145) (0.029)^2$$

Answer: 1.22×10^{-5}

K_s , solubility product/constant this is an equilibrium constant which applies to saturated solutions
a large K_s value = high solubility
a low K_s value = low solubility/insoluble

Working out the solubility product, K_s

Given $K_s = 1.81 \times 10^{-10}$ of AgCl , find the concentration of Ag^+ ions

$$K_s = [\text{Ag}^+] [\text{Cl}^-]$$

$$= x \cdot x$$

$$= x^2$$

$$\sqrt{K_s} = x$$

Answer: $1.35 \times 10^{-5} \text{ mol L}^{-1}$

Given $K_s = 7.1 \times 10^{-5}$ of CaSO_4 , find the conc. of SO_4^{2-} ions in CaSO_4

$$K_s = [\text{Ca}^{2+}] [\text{SO}_4^{2-}]$$

$$K_s = x \cdot x$$

$$\sqrt{K_s} = x$$

$$\sqrt{7.1 \times 10^{-5}} = x$$

Answer: $8.4 \times 10^{-3} \text{ mol L}^{-1}$

...more working out K_s

Given $K_s = 6.6 \times 10^{-6}$, find the concentration of Cu^{2+} ions in CuBr_2

$$K_s = [\text{Cu}^{2+}] [\text{Br}^-] [\text{Br}^-]$$

$$K_s = [\text{Cu}^{2+}] [\text{Br}^-]^2$$

$$K_s = x \cdot 2x^2$$

$$K_s = 4x^3$$

$$\sqrt[3]{K_s} = x$$

$$4$$

$$\sqrt[3]{1.65 \times 10^{-6}} = x$$

Answer: $0.0118 \text{ mol L}^{-1}$

...and more working out K_s

Given $K_s = 1.2 \times 10^{-5}$, find the concentration of Ag^+ ions in Ag_2SO_4

$$K_s = [\text{Ag}^+] [\text{Ag}^+] [\text{SO}_4^{2-}]$$

$$K_s = [\text{Ag}^+]^2 [\text{SO}_4^{2-}]$$

$$K_s = 2x^2 \cdot x$$

$$K_s = 4x^3$$

$$\sqrt[3]{K_s} = x$$

$$4$$

$$\sqrt[3]{3 \times 10^{-6}} = x$$

$$\text{So, } x = 0.0171 \text{ mol L}^{-1}$$

Answer: the concentration of Ag^+ ions is twice x , therefore $0.0342 \text{ mol L}^{-1}$

the “**common ion**” effect: precipitation can occur ie solubility decreases if an ion is added to a solution that already contains that ion

Calculate the solubility of $\text{Fe}(\text{OH})_2$ in a 0.05 mol L^{-1} solution of NaOH
 K_s of $\text{Fe}(\text{OH})_2 = 7.9 \times 10^{-16}$

$$K_s = [\text{Fe}^{2+}] [\text{OH}^-]^2$$

$$\text{assume that } [\text{OH}^-] = 0.05 \text{ mol L}^{-1}$$

$$7.9 \times 10^{-16} = [\text{Fe}^{2+}] (0.05)^2$$

$$\frac{7.9 \times 10^{-16}}{(0.05)^2} = [\text{Fe}^{2+}]$$

Answer: 3.16×10^{-13}

Calculate the solubility of AgCl in 0.1 mol L^{-1} NaCl .

$$K_s \text{ of } \text{AgCl} \text{ is } 2 \times 10^{-10}$$

$$K_s = [\text{Ag}^+] [\text{Cl}^-]$$

$$\text{assume that } [\text{Cl}^-] = 0.1 \text{ mol L}^{-1}$$

$$2 \times 10^{-10} = [\text{Ag}^+] (0.1)$$

$$\frac{2 \times 10^{-10}}{(0.1)} = [\text{Ag}^+]$$

Answer: 2×10^{-9}

If a solution is not in equilibrium the term is **Ionic product (IP)**

$\text{IP} > K_s$ a ppt will occur

$\text{IP} = K_s$ a saturated solution

$\text{IP} < K_s$ there is no ppt

Consider whether a precipitate will occur if 50 mL of 0.02 mol L^{-1} Na_2CO_3 is mixed with 50 mL of 0.05 mol L^{-1} CaCO_3 .

$$K_s (\text{CaCO}_3) = 3.4 \times 10^{-9}$$

As the two solutions are mixed together the volume doubles
so the concentration of all ions is halved

$$[\text{Ca}^{2+}] = \frac{0.05}{2} [\text{CO}_3^{2-}] = \frac{0.05}{2} + \frac{0.02}{2}$$

$$= 0.025 \quad = 0.035$$

$$\text{IP} = [\text{Ca}^{2+}] [\text{CO}_3^{2-}] = 0.025 \times 0.035 = 8.75 \times 10^{-4}$$

Answer: $\text{IP} > K_s$ so a precipitate will occur